Chapter 12 Review

Key Terms

12.1
- Bond
- Bond energy
- Ionic bonding
- Ionic compound
- Covalent bonding
- Polar covalent bond
- Electronegativity
- Dipole moment

12.3
- Lewis structure
- Duet rule
- Octet rule
- Bonding pair
- Lone pairs (unshared pairs)
- Single bond
- Double bond
- Triple bond
- Resonance

12.4
- Molecular structure (geometric structure)
- Linear structure
- Trigonal planar structure
- Tetrahedral structure
- Valence shell electron pair repulsion (VSEPR) model
- Trigonal pyramid

Key Ideas

12.1 Characteristics of Chemical Bonds

- Chemical bonds hold groups of atoms together to form molecules and ionic solids.
- Bonds are classified as:
  - Ionic: Formed when one or more electrons are transferred to form positive and negative ions
  - Covalent: Electrons are shared equally between identical atoms
    - Polar covalent bond: unequal electron sharing between different atoms
- Electronegativity is the relative ability of an atom to attract the electrons shared with another atom in a bond.
- The difference in electronegativity of the atoms forming a bond determines the polarity of that bond.

12.2 Characteristics of Ions and Ionic Compounds

- In stable compounds, atoms tend to achieve the electron configuration of the nearest noble gas atom.
- In ionic compounds:
  - Nonmetals tend to gain electrons to reach the electron configuration of the next noble gas atom.
  - Metals tend to lose electrons to reach the electron configuration of the previous noble gas atom.
- Ions group together to form compounds which are electrically neutral.

12.3 Lewis Structures

- In covalent compounds, nonmetals share electrons so that both atoms achieve noble gas configurations.
- Lewis structures represent the valence electron arrangements of the atoms in a compound.
- The rules for drawing Lewis structures recognize the importance of noble gas electron configurations.
  - Duet rule for hydrogen
  - Octet rule for most other atoms
- Some molecules have more than one valid Lewis structure, called resonance.
- Some molecules violate the octet rule for the component atoms.
  - Examples are BF_3, NO_2, and NO.

12.4 Structures of Molecules

- Molecular structure describes how the atoms in a molecule are arranged in space.
- Molecular structure can be predicted by using the valence shell electron pair repulsion (VSEPR) model.

Chapter 12 Assessment

RESEARCH LINKS
CLASSZONE.COM

All exercises with blue numbers have answers in the back of this book.

12.1 Characteristics of Chemical Bonds

A. Types of Chemical Bonds

1. Define a chemical bond.
2. What types of elements react to form *ionic* compounds? Give an example of the formation of an ionic compound from its elements.
3. What type of bonding requires the complete *transfer* of an electron from one atom to another? What type of bonding involves the sharing (either equally or unequally) of electrons between atoms?
4. Describe the type of chemical bonding that exists between the atoms in the hydrogen molecule, H_2.
5. Describe the type of chemical bonding that exists between the atoms in the hydrogen fluoride molecule, HF.

B. Electronegativity

6. What do chemists mean by the term *electronegativity*? What does its electronegativity tell us about the atom?
7. What does it mean to say that a bond is *polar*? What are the conditions that give rise to a bond's being polar?
8. For each of the following sets of elements, identify the element expected to be most electronegative and that expected to be least electronegative.

   a. K, Sc, Ca b. Br, F, At c. C, O, N

9. On the basis of the electronegativity values given in Figure 12.4, indicate whether each of the following bonds would be expected to be ionic, covalent, or polar covalent.

   a. S—S
   b. S—O
   c. S—H
   d. S—K

10. Which of the following molecules contain polar covalent bonds?

   a. phosphorus, P_4
   b. oxygen, O_2
   c. ozone, O_3
   d. hydrogen fluoride, HF

11. On the basis of the electronegativity values given in Figure 12.4, indicate which is the more polar bond in each of the following pairs.

   a. H—O or H—N
   b. H—N or H—F
   c. H—O or H—F
   d. H—O or H—Cl

12. On the basis of the electronegativity values given in Figure 12.4, indicate which bond of the following pairs has a more ionic character.

   a. Na—O or Na—N
   b. K—S or K—P
   c. Na—Cl or K—Cl
   d. Na—Cl or Mg—Cl

C. Bond Polarity and Dipole Moments

13. What is a *dipole moment*? Give four examples of molecules that possess dipole moments, and draw the direction of the dipole as shown in this section.
14. Why is the presence of a dipole moment in the water molecule so important? What are some properties of water that are determined by its polarity?
15. In each of the following diatomic molecules, which end of the molecule is positive relative to the other end?

   a. hydrogen fluoride, HF
   b. chlorine monofluoride, ClF
   c. iodine monochloride, ICl

16. For each of the following bonds, draw a figure indicating the direction of the bond dipole, including which end of the bond is positive and which is negative.

   a. P—F
   b. P—O
   c. P—C
   d. P—H

17. For each of the following bonds, draw a figure indicating the direction of the bond dipole, including which end of the bond is positive and which is negative.

   a. S—P
   b. S—O
   c. S—N
   d. S—Cl

SECTION 12.1

Characteristics of Chemical Bonds

Key Terms

- Bond
- Bond energy
- Ionic bonding
- Ionic compound
- Covalent bonding
- Polar covalent bond
- Electronegativity
- Dipole moment

Objectives

- *To learn about ionic and covalent bonds and explain how they are formed*
- *To learn about the polar covalent bond*
- *To understand the nature of bonds and their relationship to electronegativity*
- *To understand bond polarity and how it is related to molecular polarity*

Molecular bonding and structure play the central role in determining the course of chemical reactions, many of which are vital to our survival. Most reactions in biological systems are very sensitive to the structures of the participating molecules; in fact, very subtle differences in shape sometimes serve to channel the chemical reaction one way rather than another. Molecules that act as drugs must have exactly the right structure to perform their functions correctly. Structure also plays a central role in our senses of smell and taste. Substances have a particular odor because they fit into the specially shaped receptors in our nasal passages, and taste is also dependent on molecular shape.

To understand the behavior of natural materials, we must understand the nature of chemical bonding and the factors that control the structures of compounds. In this chapter, we will present various classes of compounds that illustrate the different types of bonds. We will then develop models to describe the structure and bonding that characterize the materials found in nature.

A. Types of Chemical Bonds

A water molecule

What is a chemical bond? Although there are several ways to answer this question, we will define a **bond** as a force that holds groups of two or more atoms together and makes them function as a unit. For example, in water the fundamental unit is the H—O—H molecule, which we describe as being held together by the two O—H bonds. We can obtain information about the strength of a bond by measuring the energy required to break the bond, the **bond energy**.

Atoms can interact with one another in several ways to form aggregates. We will consider specific examples to illustrate the various types of chemical bonds.

Bond (chemical bond)
The force that holds two or more atoms together and makes them function as a unit

Bond energy
The energy required to break a given chemical bond

Ionic Bonding In Chapter 8 we saw that when solid sodium chloride is dissolved in water, the resulting solution conducts electricity, a fact that convinces chemists that sodium chloride is composed of Na^+ and Cl^- ions. Thus, when sodium and chlorine react to form sodium chloride, electrons are transferred from the sodium atoms to the chlorine atoms to form Na^+ and Cl^- ions, which then aggregate to form solid sodium chloride. The resulting solid sodium chloride is a very sturdy material; it has a melting point of approximately 800 °C. See **Figure 12.1**.

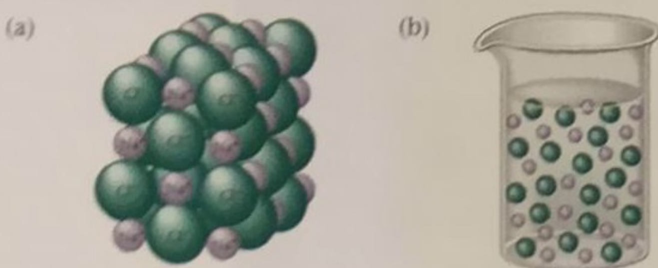
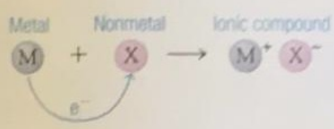


Figure 12.1
(a) Sodium chloride contains ions closely packed together in the solid. (b). When dissolved in water, these ions become free to move around

The strong bonding forces present in sodium chloride result from the attractions among the closely packed, oppositely charged ions. This is an example of **ionic bonding**. Ionic substances are formed when an atom that loses electrons relatively easily reacts with an atom that has a high affinity for electrons. In other words, an **ionic compound** results when a metal reacts with a nonmetal.

Ionic bonding
The attraction between oppositely charged ions

Ionic compound
A compound that results when a metal reacts with a nonmetal to form cations and anions



We have seen that a bonding force develops when two very different types of atoms react to form oppositely charged ions. But how does a bonding force develop between two identical atoms? Let's explore this situation by considering what happens when two hydrogen atoms are brought close together, as shown in **Figure 12.2.** When hydrogen atoms are close together, the two electrons are simultaneously attracted to both nuclei. Note in Figure 12.2b how the electron probability increases between the two nuclei, indicating that the electrons are shared by the two nuclei.

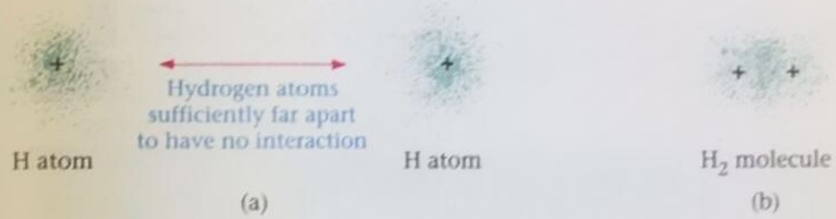


Figure 12.2
The formation of a bond between two hydrogen atoms. (a) Two separate hydrogen atoms. (b) When two hydrogen atoms come close together, the two electrons are attracted simultaneously by both nuclei. This produces the bond. Note the relatively large electron probability between the nuclei, indicating sharing of the electrons.

Covalent bonding The type of bonding we encounter in the hydrogen molecule and in many other molecules where *electrons are shared by nuclei* is called **covalent bonding**. Note that in the H_2 molecule, the electrons reside primarily in the space between the two nuclei, where they are attracted simultaneously by both protons. Although we will not go into detail about it here, the increased attractive forces in this area lead to the formation of the H_2 molecule from the two separated hydrogen atoms. When we say that a bond is formed between the hydrogen atoms, we mean that the H_2 molecule is more stable than two separated hydrogen atoms by a certain quantity of energy (the bond energy).

Covalent bonding
A type of bonding in which atoms share electrons

So far we have considered two extreme types of bonding. In ionic bonding, the participating atoms are so different that one or more electrons are transferred to form oppositely charged ions. The bonding results from the attractions between these ions. In covalent bonding, two identical atoms share electrons equally. The bonding results from the mutual attraction of the two nuclei for the shared electrons. Between these extremes are intermediate cases in which the atoms are not so different that electrons are completely transferred but are different enough so that unequal sharing

of electrons results, forming what is called a **polar covalent bond**. The hydrogen fluoride (HF) molecule contains this type of bond, which produces the charge distribution

Polar covalent bond
A covalent bond in which the electrons are not shared equally because one atom attracts the shared electrons more than the other atom

$$\underset{\delta^+}{\mathrm{H}}—\underset{\delta^-}{\mathrm{F}}$$

where δ (delta) is used to indicate a partial or fractional charge.

The most logical explanation for the development of *bond polarity* (the partial positive and negative charges on the atoms in such molecules as HF) is that the electrons in the bonds are not shared equally. For example, we can account for the polarity of the HF molecule by assuming that the fluorine atom has a stronger attraction than the hydrogen atom for the shared electrons (see **Figure 12.3**). Because bond polarity has important chemical implications, we find it useful to assign a number that indicates an atom's ability to attract shared electrons.

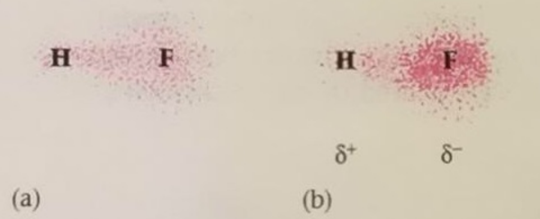


Figure 12.3
Probability representations of the electron sharing in HF. (a) What the probability map would look like if the two electrons in the H—F bond were shared equally. (b) The actual situation, where the shared pair spends more time close to the fluorine atom than to the hydrogen atom. This gives fluorine a slight excess of negative charge and the hydrogen a slight deficit of negative charge (a slight positive charge).

Active Reading Question

Provide an example of an ionic bond, an example of a polar covalent bond, and an example of a covalent bond.

B. Electronegativity

We have seen that when a metal and a nonmetal react, one or more electrons are transferred from the metal to the nonmetal to give ionic bonding. On the other hand, two identical atoms react to form a covalent bond in which electrons are shared equally. When *different* nonmetals react, a bond forms in which electrons are shared *unequally*, giving a polar covalent bond. The unequal sharing of electrons between two atoms is described by a property called **electronegativity**: *the relative ability of an atom in a molecule to attract shared electrons to itself.*

Electronegativity
The tendency of an atom in a molecule to attract shared electrons to itself

Chemists determine electronegativity values for the elements (see **Figure 12.4**) by measuring the polarities of the bonds between various atoms.

Electronegativity generally
- increases going from left to right across a period
- decreases going down a group for the representative elements.

The range of electronegativity values goes from 4.0 for fluorine to 0.7 for cesium and francium. Remember, the higher the atom's electronegativity value, the closer the shared electrons tend to be to that atom when it forms a bond.

Increasing electronegativity →

Decreasing electronegativity ↓

				H 2.1												
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La-Lu 1.0-1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac 1.1	Th 1.3	Pa 1.4	U 1.4	Np-No 1.4-1.3										

Key
- < 1.5
- 1.5–1.9
- 2.0–2.9
- 3.0–4.0

Figure 12.4
Electronegativity values for selected elements. Note that metals have relatively low electronegativity values and that nonmetals have relatively high values.

The polarity of a bond depends on the *difference* between the electronegativity values of the atoms forming the bond.

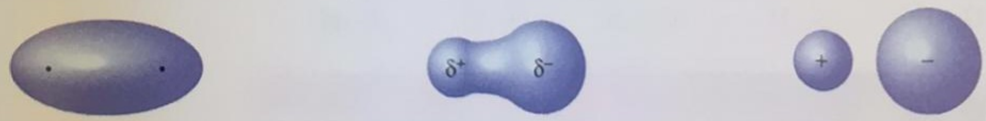
If the atoms have very similar electronegativities, the electrons are shared almost equally and the bond shows little polarity. If the atoms have very different electronegativity values, a very polar bond is formed. In extreme cases one or more electrons are actually transferred, forming ions and an ionic bond. For example, when an element from Group 1 (electronegativity values of about 0.8) reacts with an element from Group 7 (electronegativity values of about 3), ions are formed and an ionic substance results. In general, if the *difference* between the electronegativities of two elements is about 2.0 or greater, the bond is considered to be ionic.

The relationship between electronegativity and bond type is shown in **Table 12.1.** The various types of bonds are summarized in **Figure 12.5.**

CRITICAL THINKING

We use differences in electronegativity to account for certain properties of bonds.

What if all atoms had the same electronegativity values? How would bonding between atoms be affected? What are some differences we would notice?



(a) (b) (c)

Figure 12.5
The three possible types of bonds are (a) a covalent bond formed between identical atoms; (b) a polar covalent bond, with both ionic and covalent components; and (c) an ionic bond, with no electron sharing.

Active Reading Question

In general, how do electronegativity trends compare to trends of atomic size?

Table 12.1

The Relationship Between Electronegativity and Bond Type

Electronegativity Difference Between the Bonding Atoms	Bond Type	Covalent Character	Ionic Character
Zero	Covalent	Increases ↑	Increases ↓
↓	↓		
Intermediate	Polar covalent		
↓	↓		
Large	Ionic		

EXAMPLE 12.1

Using Electronegativity to Determine Bond Polarity

Using the electronegativity values given in Figure 12.4, arrange the following bonds in order of increasing polarity: H—H, O—H, Cl—H, S—H, and F—H.

Solution

The polarity of the bond increases as the difference in electronegativity increases. From the electronegativity values in Figure 12.4, the following variation in bond polarity is expected (the electronegativity value appears in parentheses below each element).

Bond	Electronegativity Value	Difference in Electronegativity Values	Bond Type	Polarity
H—H	(2.1) (2.1)	2.1 − 2.1 = 0	Covalent	Increasing ↓
S—H	(2.5) (2.1)	2.5 − 2.1 = 0.4	Polar covalent	
Cl—H	(3.0) (2.1)	3.0 − 2.1 = 0.9	Polar covalent	
O—H	(3.5) (2.1)	3.5 − 2.1 = 1.4	Polar covalent	
F—H	(4.0) (2.1)	4.0 − 2.1 = 1.9	Polar covalent	

Therefore, in order of increasing polarity, we have

H—H S—H Cl—H O—H F—H

Least polar → Most polar

Practice Problem • Exercise 12.1

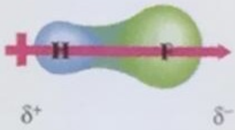
For each of the following pairs of bonds, choose the bond that is more polar.

a. H—P, H—C
b. O—F, O—I
c. N—O, S—O
d. N—H, Si—H

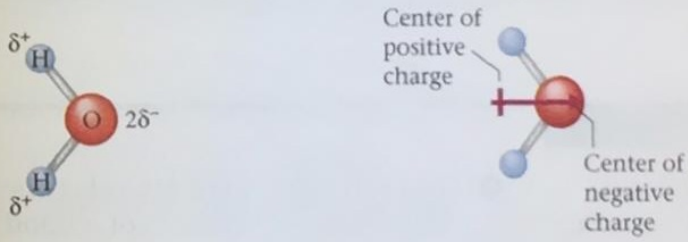
C. Bond Polarity and Dipole Moments

We have seen that hydrogen fluoride has a positive end and a negative end. A molecule such as HF that has a center of positive charge and a center of negative charge is said to have a **dipole moment**. The dipolar character of a molecule is often represented by an arrow. This arrow points toward the negative charge center, and its tail indicates the positive center of charge:

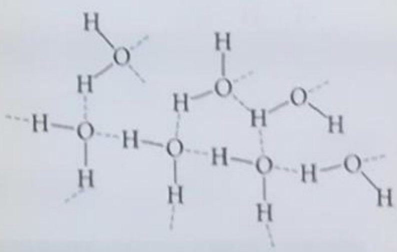
Dipole moment
A property of a molecule in which the charge distribution can be represented by a center of positive charge and a center of negative charge



Any diatomic (two-atom) molecule that has a polar bond has a dipole moment. Some polyatomic (more than two atoms) molecules also have dipole moments. For example, because the oxygen atom in the water molecule has a greater electronegativity than the hydrogen atoms, the electrons are not shared equally. This results in a charge distribution that causes the molecule to behave as though it had two centers of charge—one positive and one negative. So the water molecule has a dipole moment.



The fact that the water molecule is polar (has a dipole moment) has a profound effect on its properties. In fact, it is not overly dramatic to state that the polarity of the water molecule is crucial to life as we know it on earth. Because water molecules are polar, they can surround and attract both positive and negative ions (see **Figure 12.6**). These attractions allow ionic materials to dissolve in water. Also, the polarity of water molecules causes them to attract each other strongly.



This means that much energy is required to change water from a liquid to a gas (the molecules must be separated from each other to undergo this change of state). Therefore, it is the polarity of the water molecule that causes water to remain a liquid at the temperatures on the earth's surface. If it were nonpolar, water would be a gas and the oceans would be empty.

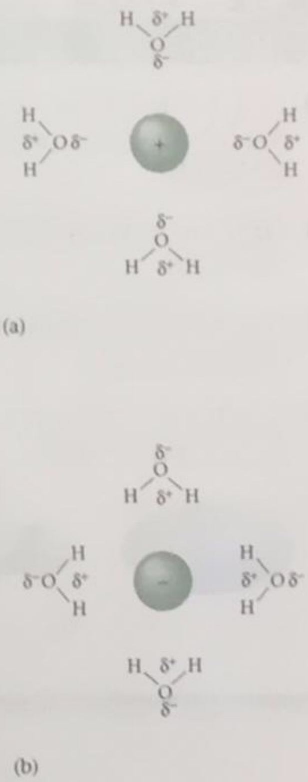


Figure 12.6
(a) Polar water molecules are strongly attracted to positive ions by their negative ends. (b) They are also strongly attracted to negative ions by their positive ends.